Phase II - Molecules
Molecules (chpt 2)

1) Which group in the periodic table contains only nonmetals?
A) 1A B) 6A C) 2B D) 2A E) 8A

2) The element _____ is the most similar to strontium in chemical and physical properties.
A) Li B) At C) Rb D) Ba E) Cs

3) Lithium is a _____ and magnesium is a _____.
A) nonmetal, metal B) nonmetal, nonmetal C) metal, metal D) metal, metalloid E) metalloid, metalloid

4) Calcium is a _____ and silver is a _____.
A) nonmetal, metal B) metal, metal C) metalloid, metal D) metal, metalloid E) nonmetal, metalloid

5) When a metal and a nonmetal react, the _____ tends to lose electrons and the _____ tends to gain electrons.
A) metal, metal B) nonmetal, nonmetal C) metal, nonmetal D) nonmetal, metal E) None of the above

6) The empirical formula of a compound that contains 12 carbon atoms, 14 hydrogen atoms, and 6 oxygen atoms is _____.

7) The formula of a salt is XCl_2. The X-ion in this salt has 28 electrons. The metal X is _____.

8) Predict the empirical formula of the ionic compound that forms from magnesium and fluorine.

9) Name the following: a. SO b. H_2CO_3 (an acid) c. H_2SO_3 (an acid) d. $CrCl_3$. e. $Ni(CN)_2$

11) The formula of ammonium carbonate is _____.

12) The formula for aluminum hydroxide is _____.

13. What is the formula for chromium(III) oxide?

14. Predict the empirical formulas of the ionic compounds formed from the following pairs of elements. Name each compound.
a. K and N b. Ga and O c. Rb and Cl d. Ba and S

Stoichiometry (chpt 3)

15) How many grams of hydrogen are in 46 g of CH_4O?

16) A 2.25-g sample of magnesium nitrate, $Mg(NO_3)_2$, contains _____ mol of this compound.

17) A compound that is composed of carbon, hydrogen, and oxygen contains 70.6% C, 5.9% H, and 23.5% O by mass. The molecular weight of the compound is 136 amu. What is the molecular formula?

18) The formula weight of potassium dichromate $(K_2Cr_2O_7)$ is _____ amu.

19) The mass % of Al in aluminum sulfate $(Al_2(SO_4)_3)$ is _____.

21. Aspartame is an artificial sweetener that is 160 times sweeter than sucrose (table sugar) when dissolved in water. It is marketed by G.D. Searle as Nutra-Sweet. The molecular formula of aspartame is $C_{14}H_{18}N_2O_5$.
a. Calculate the molecular weight of aspartame.
b. How many moles of molecules are in 10.0g of aspartame?
c. What is the mass in grams of 1.56 mol of aspartame?
d. How many molecules are in 5.0 mg of aspartame?
e. How many atoms of nitrogen are in 1.2 g of aspartame?
f. What is the mass in grams of 1.0×10^9 molecules of aspartame?

22. Calculate the molecular weight of each of the following substances:
a. $CuCl_2 \cdot 6H_2O$ (The $\cdot 6H_2O$ indicates the presence of six water molecules.)

b. $NaBrO_3$ c. H_3PO_4 d.$CaCO_3$

23. Express the composition of each compound as the mass percents of its elements.
a. formaldehyde, CH_2O b. glucose, $C_6H_{12}O_6$ c. acetic acid, $HC_2H_3O_2$

24. What is the empirical formula for each of the following compounds (where the common names are given)?
a. vitamin C, $C_6H_8O_6$ d. phosphorus (V) oxide, P_2O_{10}
b. benzene, C_6H_6 e. glucose, $C_6H_{12}O_6$
c. acetylene, C_2H_2 f. acetic acid, $HC_2H_3O_2$

25. A compound contains only carbon, hydrogen, and oxygen is 48.38% C and 8.12% H by mass. What is the empirical formula of this substance?

26. A compound contains only carbon, hydrogen, nitrogen, and oxygen. Combustion of 0.157 g of the compound produced 0.213 g of CO_2 and 0.0310 g of H_2O. In another experiment, it is found that 0.103 g of the compound procedures 0.0230 go of NH_3. What is the empirical formula of the compound? Hint: Combustion involves reacting with excess O_2. Assume that all of the carbon ends up in CO_2 and all of the hydrogen ends up in H_2O.

Bonding (chpt 8)
27) There are _____ paired and _____ unpaired electrons in the ground state e- configuration of a phosphorus atom.

28) In the Lewis symbol for a fluorine atom, there are _____ paired and _____ unpaired electrons.

29) Which of the following would have to lose two electrons in order to achieve a noble gas electron configuration _____? O Sr Na Se Br

31) Which of the following would have to gain two electrons in order to achieve a noble gas electron configuration _____? O Sr Na Se Br

32) For a given arrangement of ions, the lattice energy increases as ionic radius _____ and as ionic charge _____.

33) Elements from opposite sides of the periodic table tend to form _____.
A) covalent compounds B) ionic compounds C) compounds that are gaseous at room temperature
D) homonuclear diatomic compounds E) covalent compounds that are gaseous at room temperature

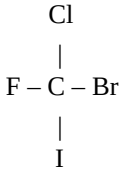
34) How many single covalent bonds must a silicon atom form to have a complete octet in its valence shell?

35) A double bond consists of _____ pairs of electrons shared between two atoms.

36) What is the maximum number of double bonds that a hydrogen atom can form?

37) What is the maximum number of double bonds that a carbon atom can form?

38) In the molecule below, which atom has the largest partial negative charge _____?



39) Given the electronegativities below, a bond between which 2 atoms forms a covalent single bond that is most polar?

Element:	H	C	N	O
Electronegativity:	2.1	2.5	3.0	3.5

41) A nonpolar bond will form between two _____ atoms of _____ electronegativity.
A) different, opposite B) identical, different C) different, different D) similar, different E) identical, equal

42. (a) Define the term lattice energy. (b) Which factors govern the magnitude of the lattice energy of an ionic compound?

43. (a) Does the lattice energy of an ionic solid increase or decrease (i) as the charges of the ions increase, (ii) as the sizes of the ions increase? (b) Using a periodic table, arrange the following substances according to their expected lattice energies, listing them from lowest lattice energy to the highest: LiCl, NaBr, RbBr, MgO.

44. Explain the following trends in lattice energy: (a) Mg) > $MgCl_2$; (b) NaCl > RbBr > CsBr; (c) BaO > KF.

45. Which of the following bonds are polar: (a) B-F, (b) Cl –Cl, (c) Se –O, (d) H – I? Which is the more electronegative atom in each polar bond?

46. Arrange the bonds in each of the following sets in order of increasing polarity:
(a) C – F, O – F, Be – F; (b) O – Cl, S – Br, C – P; (c) C – S, B –F, N – O.

47. Give the name or chemical formula, as appropriate, for each of the following substances, and in each case predict whether the bonding is better described by the ionic-bonding or covalent-bonding model:
(a) manganese (III) fluoride, (b) chromium (VI) oxide, (c) arsenic (V) bromide, (d) SF_4, (e) $MoCl_4$, (f) $ScCl_3$.

48. Without using a table of electronegativities, predict the order of increasing electronegativity in each of the following groups of elements:
a. C, N, O b. S, Se, Cl c. Si, Ge, Sn d. Tl, S, Ge

49. Predict which bond in each of the following groups will be the most polar:
a. C–F, Si–F, Ge–F b. P–Cl, S–Cl c. S–F, S–Cl, S–Br d. Ti–Cl, Si–Cl, Ge–Cl
e. C–H, Si–H, Sn–H f. Al–Br, Ga–Br, In–Br, Tl–Br

51. Define the term isoelectronic.

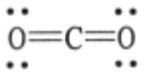
52. When comparing sizes of ions of elements in the same period in the same periodic table, why is it advantageous to compare isoelectronic species?

53. Which compound in each of the following pairs of ionic substances has the greatest lattice energy? Justify your answers.
a. NaCL, KCl b. LiF, LiCl c. $Mg(OH)_2$, MgO d. $Fe(OH)_2$, $Fe(OH)_3$
Give three ions that are isoelectronic with the neutral krypton atom.

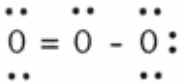
54) The Lewis structure of AsH_3 shows _____ nonbonding electron pair(s) on As.

55) The Lewis structure of PF_3 shows that the central phosphorus atom has _____ nonbonding and _____ bonding electron pairs.

56) The formal charge on carbon in the molecule below is _____.



57) In the resonance form of ozone shown below, the formal charge on the central oxygen atom is _____.



58) How many equivalent resonance forms can be drawn for SO_2 without expanding octet on the S atom (S is the central atom)?

59) Lattice energy is _____.
A) the energy required to convert a mole of ionic solid into its constituent ions in the gas phase
B) the energy given off when gaseous ions combine to form one mole of an ionic solid
C) the energy required to produce one mole of an ionic compound from its constituent elements in their standard states
D) the sum of ionization energies of the components in an ionic solid
E) the sum of electron affinities of the components in an ionic solid

61) The type of compound that is most likely to contain a covalent bond is _____.
A) one that is composed of a metal from the far left of the periodic table and a nonmetal from the far right of the periodic table

B) a solid metal
C) one that is composed of only nonmetals
D) held together by the electrostatic forces between oppositely charged ions
E) There is no general rule to predict covalency in bonds.

62) In which of the molecules below is the carbon-carbon distance the shortest?
A) $H_2C{=}CH_2$ B) $H{-}C{\equiv}C{-}H$ C) $H_3C{-}CH_3$ D) $H_2C{=}C{=}CH_2$ E) $H_3C{-}CH_2{-}CH_3$

63) Which of the following has the bonds correctly arranged in order of increasing polarity?
A) Be–F, Mg–F, N–F, O–F B) O–F, N–F, Be–F, Mg–F C) O–F, Be–F, Mg–F, N–F
D) N–F, Be–F, Mg–F, O–F E) Mg–F, Be–F, N–F, O–F

64) Resonance structures differ by _____.
A) number and placement of electrons B) number of electrons only C) placement of atoms only
D) number of atoms only E) placement of electrons only

65) For resonance forms of a molecule or ion, _____.
A) one always corresponds to the observed structure
B) all the resonance structures are observed in various proportions
C) the observed structure is an average of the resonance forms
D) the same atoms need not be bonded to each other in all resonance forms
E) there cannot be more than two resonance structures for a given species

66) As the number of covalent bonds between two atoms increases, the distance between the atoms _____ and the strength of the bond between them _____.
A) increases, increases B) decreases, decreases C) increases, decreases D) decreases, increases E) is unpredictable

67) Draw the most stable Lewis structure (or 'explain multiple stable resonance structures) of the following:
In each case give the bond order of the 'actual' structure.

(i) NO_2^- (ii) SO_3^{2-} (iii) SO_4^{2-} (iv) Br_3^- (v) SO_2

68. Tell the difference between the following pairs of terms:
a. electronegativity and electron affinity
b. covalent bond and polar bond
c. polar covalent bond and ionic bond

69. Write a Lewis structure for each of the following molecules and ions. In each case the first atom listed is the central atom.
a. $POCl_3$, PO_3^{2-}, PCl_2^-

71. Write Lewis structures for the following. Show all resonance structures where applicable.
a. NO_2^-, HNO_2, NO_3^-, HNO_3
b. SO_4^{2-}, HSO_4^-, H_2SO_4
c. CN^-, HCN
d. OCN^-, SCN^-, N_3^-
e. C_2N_2 (atomic arrangement is NCCN)

72. Write Lewis structures for the following molecules, which have central atoms that do not obey the octet rule: PF_5, BrF_3, $Be(CH_3)_2$, BCl_3, $XeOF_4$ (Xe is the central atom), XeF_6, SeF_4.

73. Predict the molecular structure of each of the following I_3^-, ClF_3, IF_4^+, and SF_5^+.

74. Predict the molecular structures of BrF_5, KrF_4, and IF_6^+.

75. What two requirements must be satisfied for a molecule to be polar?

76. Write a Lewis structure and predict the molecular structure and polarity of each of the following sulfur fluorides: SF_2, SF_4, SF_6, and S_2F_4 (exists as $F_3S - SF$)
Predict the F – S - F bond angles in each molecule.

77. There are three possible isomers of $PF_3(CH_3)_2$, where P is the central atom. Draw them and describe how measurements of

dipole moments might be used to distinguish among them.

78. Draw Lewis structures for the following: (a) SiH_4, (b) CO, (c) SF_2, (d) H_2SO_4 (H is bonded to O), (e) ClO_2 , (f) NH_2OH.

79. (a) Write a Lewis structure for the phosphorus trifluoride molecule, PF_3. Is the octet rule satisfied for all the atoms in your structure? (b) Determine the oxidation numbers of the P and F atoms. (c) Determine the formal charges of the P and F atoms. (d) Is the oxidation number for the P atom the same as its formal charge? Explain why or why not.

81. For each of the following molecules or ions of sulfur and oxygen, write a single Lewis structure that obeys the octet rule, and calculate the oxidation numbers and formal charges on all the atoms: (a) SO_2, (b) SO_3, (c) SO_3^{2-}, (d) SO_4^{2-}.

82. Predict the ordering of the C –O bond lengths in CO, CO_2, and CO_3^{2-}.

83. Draw the Lewis structures for each of the following ions or molecules. Identify those that do not obey the octet rule, and explain why they do not. (a) SO_3^{2-}, (b) AlH_3, (c) N_3^-, (d) CH_2Cl_2, (e) SbF_5.

84. (a) Describe the molecule chlorine dioxide, ClO_2, using three possible resonance structures. (b) Do any of these resonance structures satisfy the octet rule for every atom in the molecule? Why or why not? (c) Using formal charges, select the resonance structure(s) that is (are) most important.

85. Three possible Lewis structures can be written for nitrous oxide (N_2O):

N=N=0←→: N≡N–O ←→N–N≡O :

N-N	1.67 À	N=O	1.15À
N=N	1.20 À	N—O	1.47 À
N≡N	1.10 À		

86. Calculate the formal charge on the indicated atom in each of the following molecules or ions: (a) the central oxygen atom in O_3, (b) phosphorus in PF_6^-, (c) nitrogen in NO_2, (d) iondine in ICl_3, (e) Chlorine in $HClO_4$ (hydrogen is bonded to O).

Molecular Structure (chpt 9)

87) For a molecule with the formula AB_2 the molecular shape could be _____.

88) According to VSEPR theory, if there are five electron domains in the valence shell of an atom, they will be arranged in a(n) _____ geometry.

89) According to VSEPR theory, if there are four electron domains in the valence shell of an atom, they will be arranged in a(n) _____ geometry.

91) The electron-domain geometry and molecular geometry of iodine trichloride are _____ and _____, respectively.

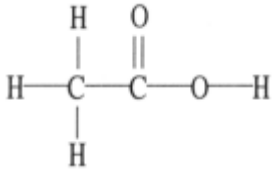
92. (a) How does a polar molecule differ from a nonpolar one? (b) Atoms X and Y have different electronegativities. Will the diatomic molecule X – Y necessarily be polar? Explain. (c) What factors affect the size of the dipole moment of a diatomic molecule?

93) For each of the following compounds, draw the most stable Lewis structures, give electron domain geometry, molecular geometry, approximate bond angles, hybridization on central atom, and indicate if there is an **overall** dipole moment.

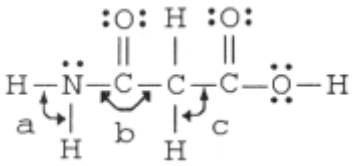
a. H_3O^+ b.IF_3 c.KrF_2 d.AsF_5 e.CS_2 f.$PHCl_2$ g.IF_5 h.PF_2^+ h.BF_2^- i.$SiCl_2F_2$ j.SO_2 k. XeF_4 l.BrO_3^- m.ICl_4^-
n. CF_4 o. TeF_4 p. $XeOF_4$ q. NF_3 r.AsF_5 s. $XeOF_2$ t. OF_2 u.KrF_2 v. XeO_4 w. BF_3 x.KrF_4 y. BeH_2 z. SeF_6

94) The Cl–C–Cl bond angle in the CCl_2O molecule (C is the central atom) is _____.

95) The molecular geometry of the 2 carbon atoms in the molecule below are _____ and _____.



96) The bond angles marked a, b, and c in the molecule below are about _____, _____, and _____, respectively.



97. (a) Methane (CH_4) and the perchlorate ion (ClO_4) are both described as tetrahedral. What does this indicate about their bond angles? (b) The NH_3 molecule is trigonal pyramidal. How many parameters need to be specified to define its geometry completely?

98. (a) How does one determine the number of electron domains in a molecule or ion? (b) What is the difference between a bonding electron and a non bonding electron domain?

99. Indicate the number of electron domains about a central atom, given the following angles between them: (a) 120°, (b) 180°, (c) 109.5°, (d) 90°.

101. What is the difference between the electron-domain geometry and the molecular geometry of a molecule? Use the water molecule as an example in your discussion.

102. Draw the Lewis structure for each of the following molecules or ions, and predict their electron-domain and molecular geometries: (a) PF_3, (b) CH_3^+, (c) BrF_3, (d) ClO_4^-, (e) XeF_2, (f) BrO_2^-.

103. Give the electron-domain and molecular geometries for the following molecules and ions: (a) HCN, (b) SO_3^{2-}, (c) SF_4, (d) PF_6^-, (e) NH_3Cl^+, (f) N_3^-.

104. The three species NH_2^-, NH_3, and NH_4^+ have H – N – H bond angles of 105°, 107°, and 109°, respectively. Explain this variation in bond angles.

105. Predict the trend in the F(axial) – A – F(equatorial) bond angle in the following AF_n molecules: PF_5, SF_4, and ClF_3.

106. (a) Explain why BrF_4^- is square planar, whereas BF_4^- is tetrahedral. (b) In which of these molecules, CF_4 or SF_4, do you think the actual bond angle is closest to the ideal angle predicted by the VSEPR model? Explain briefly.

107. (a) What conditions must be met if a molecule with polar bonds in nonpolar? (b) What geometries will give nonpolar molecules for AB_2, AB_3, and AB_4 geometries?

108. (a) Draw Lewis structures for methane (CH_4) and formaldehyde (H_2CO). (b) What is the hybridization at the carbon atom in CH4 and H_2CO? (c) The carbon atom in CH_4 cannot participate in multiple bonding, whereas that in H_2CO can. Explain this observation by using the hybridization at the carbon atom.

109. The molecules SiF_4, SF_4, and XeF_4 all have molecular formulas of the type AF_4, but the molecules have different molecular geometries. Predict the shape of each molecule, and explain why the shapes differ.

111. The PF_3 molecule has a dipole moment of 1.03 D, but BF_3 has no dipole moment. How can you explain the difference?

112. Sulfur tetrafluoride reacts slowly with O_2 to form sulfur tetrafluoride monoxide (OSF_4) The O atom and the 4 F atoms are all bonded to the central S atom.
a. Write and balance the chemical equation.
b. Write a Lewis structure of the product in which all formal charges are 0.
c. What is the EDG and Mg for the product?

113. The phosphorus trihalides (PX_3) show the following variation in bond angle X-P-X: PF_3 96.3 °, PCl_3 100.3°, PBr_3, 101.0°, PI_3 102°. The trend is generally attributed to the electronegativity of the halogen.
a. Assuming that all the electron domains exhibit the same repulsion, what value of the X-P-X angle is predicted by VSEPR?
b. What is the general trend in the X-P-X angle as the electronegativity increases?
c. Using the VSEPR model, explain the observed trend in X-P-X angle as electronegativity of X changes.
d. Based on your answer to part c, predict the structure of $PBrCl_4$.

114) The electron-domain geometry of a sulfur-centered compound is trigonal bipyramidal. The hybridization of the central sulfur nitrogen atom is _____.

115) The hybridization of orbitals on the central atom in a molecule is sp^2. The electron-domain geometry about this central atom is _____.

116) There are _____ σ and _____ π bonds in the H–C≡C–H molecule.

117. There are _____ σ and _____ π bonds in the $H_2C{=}C{=}CH_2$ molecule.

118. The hybridizations of the carbon and oxygen atoms in carbon monoxide are _____ and _____, respectively.

119. Indicate the hybridization of the central atom in (a) BCl_3, (b) $AlCl_4^-$, (c) CS_2, (d) KrF_2, (e) PF_6^-.

121. What set of hybrid orbitals is used by the central atom in (a) $SiCl_4$ (b) HCN (c) SO_3 (d) ICl_2^- (e) BrF_4^-?

122) A typical double bond _____.
A) is stronger and shorter than a single bond B) consists of one σ bond and one π bond C) imparts rigidity to a molecule
D) consists of two shared electron pairs E) All of the above answers are correct.

123) A typical triple bond _____.
A) consists of one σ bond and two π bonds B) consists of three shared electrons c) consists of two σ bonds and one π bond
D) consists of six shared electron pairs E) is longer than a single bond

124) The carbon-carbon σ bond in ethylene, CH_2CH_2, results from the overlap of _____.
A) sp hybrid orbitals B) sp^3 hybrid orbitals C) sp^2 hybrid orbitals D) s atomic orbitals E) p atomic orbitals

125) The π bond in ethylene, CH_2CH_2, results from the overlap of _____.
A) sp^3 hybrid orbitals B) s atomic orbitals C) sp hybrid orbitals D) sp^2 hybrid orbitals E) p atomic orbitals

126. According to valence bond theory, which orbitals on bromine atoms overlap in the formation of the bond in Br_2 ?

127) The N–N bond in HNNH consists of _____.
A) one σ bond and one π bond B) one σ bond and two π bonds C) two σ bonds and one π bond
D) two σ bonds and two π bonds E) one σ bond and no π bonds

128) An antibonding π orbital contains a maximum of _____ electrons.

129) According to MO theory, overlap of two s atomic orbitals produces _____.
A) one bonding molecular orbital and one hybrid orbital
B) two bonding molecular orbitals
C) two bonding molecular orbitals and two antibonding molecular orbitals
D) two bonding molecular orbitals and one antibonding molecular orbital
E) one bonding molecular orbital and one antibonding molecular orbital

131) A molecular orbital can accommodate a maximum of _____ electron(s).

132) An antibonding MO _____ the corresponding bonding MO.
A) is always lower in energy than B) can accommodate more electrons than C) can accommodate fewer electrons than
D) is always higher in energy than E) is always degenerate with